CHM129
Bohr Model of the Hydrogen Atom

1. Calculate the **energy, in J,** of the photons emitted for the following transitions (pay attention to the sign of your answer):
   a. from n=3 to n=2
   b. from n=4 to n=2
   c. from n=5 to n=2
   d. from n=6 to n=2

2. Find the **wavelength** and **frequency** of radiation that correspond to each transition. What color of light do these correspond to?

3. Calculate the **energy (in kJ/mol)**.

$$E_n = -2.18\times10^{-18}\,J\left(\frac{1}{n^2}\right) \qquad where\ n = 1,2,3\ldots$$

$$\Delta E = E_{final} - E_{initial} = h\nu = \frac{hc}{\lambda}$$

The difference (ΔE) in energy is the final energy-initial energy.

Constants: $h=6.626 \times 10^{-34}$ J.s $\quad$ $c= 3.00 \times 10^{8}$ m/s
$N_A=6.022 \times 10^{23}$ mol^{-1}

1. (a) $n=3 \longrightarrow n=2$

$$\Delta E = -2.18 \times 10^{-18} J\left(\frac{1}{2^2} - \frac{1}{3^2}\right) = -3.03 \times 10^{-19} J$$

(b) $n=4 \longrightarrow n=2$

$$\Delta E = -2.18 \times 10^{-18} J\left(\frac{1}{2^2} - \frac{1}{4^2}\right) = -4.09 \times 10^{-19} J$$

(C) $n=5 \longrightarrow n=2$

$$\Delta E = -2.18 \times 10^{-18} J\left(\frac{1}{2^2} - \frac{1}{5^2}\right) = -4.58 \times 10^{-19} J$$

(d) $n=6 \longrightarrow n=2$

$$\Delta E = -2.18 \times 10^{-18} J\left(\frac{1}{2^2} - \frac{1}{6^2}\right) = -4.84 \times 10^{-19} J$$

2. (a) $\nu = \frac{\Delta E}{h} = \frac{3.03 \times 10^{-19}\ J}{6.626 \times 10^{-34}\ J \cdot s} = 4.57 \times 10^{14}\ s^{-1}$

$\lambda = \frac{hc}{\Delta E} = \frac{(6.626 \times 10^{-34}\ J \cdot s)(3.00 \times 10^{8}\ m/s)}{3.03 \times 10^{-19}\ J} = 6.57 \times 10^{-7}\ m \left(\frac{1\ nm}{10^{-9}\ m}\right) = \underline{657\ nm}$

Red

(b) $\nu = \frac{4.09 \times 10^{-19}\ J}{6.626 \times 10^{-34}\ J \cdot s} = 6.17 \times 10^{14}\ s^{-1}$

$\lambda = \frac{(6.626 \times 10^{-34}\ J \cdot s)(3.00 \times 10^{8}\ m/s)}{4.09 \times 10^{-19}\ J} = 4.86 \times 10^{-7}\ m \left(\frac{1\ nm}{10^{-9}\ m}\right) = \underline{486\ nm}$

Blue-Green

(c) $\nu = \frac{4.58 \times 10^{-19}\ J}{6.626 \times 10^{-34}\ J \cdot s} = 6.91 \times 10^{14}\ s^{-1}$

$\lambda = \frac{(6.626 \times 10^{-34}\ J \cdot s)(3.00 \times 10^{8}\ m/s)}{4.58 \times 10^{-19}\ J} = 4.34 \times 10^{-7}\ m \left(\frac{1\ nm}{10^{-9}\ m}\right) = \underline{434\ nm}$

Violet

(d) $\nu = \frac{4.84 \times 10^{-19}\ J}{6.626 \times 10^{-34}\ J \cdot s} = 7.30 \times 10^{14}\ s^{-1}$

$\lambda = \frac{(6.626 \times 10^{-34}\ J \cdot s)(3.00 \times 10^{8}\ m/s)}{4.84 \times 10^{-19}\ J} = 4.10 \times 10^{-7}\ m \left(\frac{1\ nm}{10^{-9}\ m}\right) = \underline{410\ nm}$

Violet

3. (a) $3.03 \times 10^{-19}\ J \left(\frac{1\ kJ}{1000\ J}\right)\left(\frac{6.022 \times 10^{23}}{1\ mol}\right) = 182\ kJ/mol$

(b) $4.09 \times 10^{-19}\ J \left(\frac{1\ kJ}{1000\ J}\right)\left(\frac{6.022 \times 10^{23}}{1\ mol}\right) = 246\ kJ/mol$

(c) $4.58 \times 10^{-19}\ J \left(\frac{1\ kJ}{1000\ J}\right)\left(\frac{6.022 \times 10^{23}}{1\ mol}\right) = 276\ kJ/mol$

(d) $4.84 \times 10^{-19}\ J \left(\frac{1\ kJ}{1000\ J}\right)\left(\frac{6.022 \times 10^{23}}{1\ mol}\right) = 291\ kJ/mol$